Geel 2000 Language Schools



First Term

2024 - 2025

Chapter one

Structure of the atom

Atomic structure

1- Greek philosophers:

Any piece of matter can be divided into smaller parts and each part can be subdivided into smaller parts which can't be divided this part is called Atom.

2- Aristotle

A-He refused the idea of Greek philosophers about the atom.

b- He supposed that all matters composed of 4 constituents which are (water, air , dust and fire and postulated that the cheap metals such as iron or copper can be changed into precious ones like gold by changing the percentage of four constituents.

3- Boyl (Irish 1661)

i- He refused the Aristotle concept.

ii-He was the 1st scientist to define the

element as pure simple substance which cannot be analvsed into simple one by normal chemical methods.

4) Dulton s atom (Englsh 1803)

He supposed that:

1- The substance consists of very small particles called atoms.

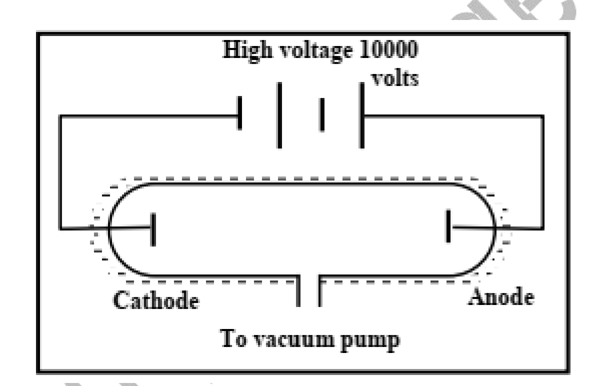
2- Every element consists of very small dense atoms which can'tbe divided.

3- Atoms f the same element are identical.

4- Atom of different are different.

N.B By the 19 th century
Scientists had accepted the idea that elements consists of atoms but they
knew very little about those atoms.

CATHODE- RAVS EXPERIMENT (DISCOVERY OF THE ELECTRON)



a-All gases under normal conditions of pressure and temp (76 cm. hg. 25c) don't conduct electricity.

b-If a glass tube evacuated from the gas to decrease its pressureto reach
0.01.1 ------------ 0.001 m.mhg
the gas will conduct electric current.

c-If the potential difference between the tow poles increases up 10.000 Volts a Flow of invisible rays are emitted from the cathode causing glowing to the wall of tube behind the anode and called cathode ray

PROPERTIES OF CATHODE RAYS

1)Consists of tiny particles have mass and velocity.

2)Transfers in straight lines glowing the glass facing the cathode.3) Have negative charge.

4)Have a thermal effect.

5)Affected by electric and Magnetic field.

6)Cathode rays don t change by changing either cathode material or type of the gas which proves that cathode rays take part in the structure of all substances.

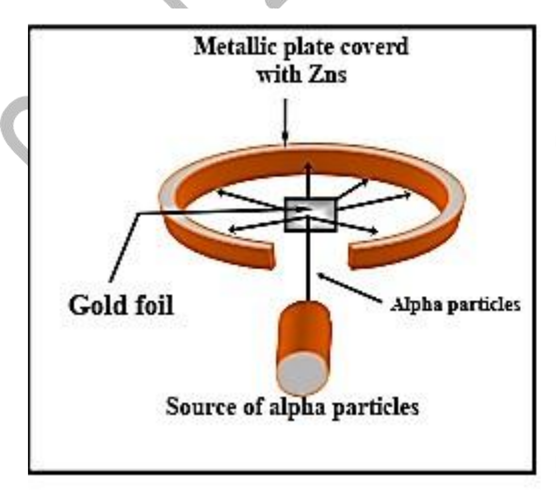
5 - Thomson s atom 1897

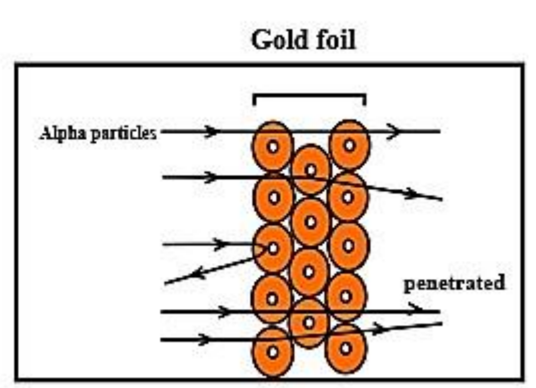
He conclude from the last experiment that
i- The atom is a homogenous sphere of positive electricity.

ii- Inside it there are negative electrons enough to make it electrically neutral.

RUTHER FORD S EXPERIMENT

In 1911 Geiger and Marsden performed a famous exp According to suggestion of Ruther ford by the following apparatus.





1 *– He allowed alpha particles to hit a metallic plate lined with Zinc suphide (glows when hits with alpha rays)*

2 *– On placing a gold foil in the front of alpha rays he concluded the following from the following observation.*

Observation	Result
1- Most of alpha particles appeared in the same position before putting gold foil.	**1- Most of the atom is a space not solid as explained by Dalton and Thomson.**
2- A very small percentage of alpha particles reflected back to appears as flashes on other side of sheet.	**2- The atom has very small part with very small volume but high density.**
3- Some flashes appeared on the sides of 1st site.	**3-The dense part of the atom which concentrate in it most mass have same charge of alpha particle (+ve) which called nucleus of the atom.**

Structure of the atom

MASS NO:

Sum of protons and neutrons inside the nucleus.

ISOTOPES:

Are different forms of atoms of the same element which have same atomic no. but different mass no.

Ex: hydrogen has 3 isotopes: ……….

Rutherford:

Rutherford was the 1st scientist who stated the concept of the atomic structure.

1- THE ATOM:

Although it has very small size but it has a complicated structure that resembles the solar system in which electrons revolve around the central nucleus in orbits as planets revolve around the sun.

2-The Nucleus:

Is much smaller than the atom. Located in the centre of the atom with (+ve) charge. There is a big space between the nucleus and orbits of electrons, so most of the atom is a space. Most mass of the atom is concentrated in the nucleus as mass of e is very small and can be neglected.

3- ELECTRONS:

1. Have negligible mass compared to that of the nucleus.
2. No of electrons (-ve) equals no of ptotons (+ve) so the atom is electrically neutral.

3. Electrons revolve around the nucleus in a fixed orbit aselectrons are affected by two forces equal in strengthbut in opposite direction, which are :

1-Force of attraction of the nucleus to electrons.

2-Centrifugal force due to velocity of electron around the nucleus.

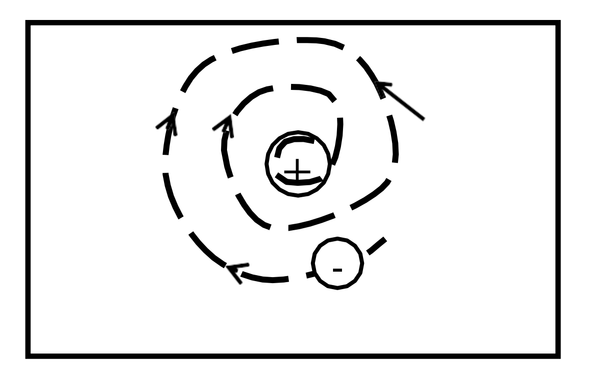
Give reason: Electrons are not attracted to thenucleus.
Explain: Structure of the atom in the view ofrutherford.

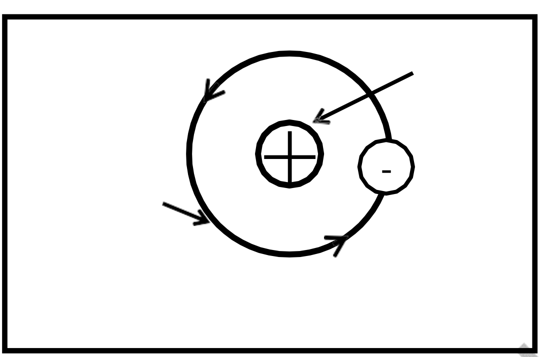
Objections on Rutherford's atomic model (Maxwell's theory):

Rutherford's concept was contradicted by maxwell's theory (Which was based on laws of newton mechanics and concerned with the movement of relatively large bodies).

1 Which states that: "When an electrecally charged body moves in orbit, it will lose its energy gradually by emission of radiation causing gradual decrease in orbit radius".

By applying this theory on electron movement in Rutherford's atom, we would expect that electrons are in a state of continuous emission of radiation, so the atomic radius will decrease and electrons move in a spiral orbit until they hit the nucleus.





GIVE REASON: Contradiction between the classical mechanical laws and rutherford.

Bohr's Atomic Model

Bohr's postulates:

1- A positively charged nucleus exists in thecenter of the atom.
2- Atom is electrically neutral as no of p^+ sequals to no of e's.

3-Electrons revolve around the nucleus in orbits due to centrifugal and attraction forces.

4-Electrons orbit the nucleus in a rapid movement without gaining or losing energy.

5-Electrons orbit the nucleus only in a definite allowed energylevels, so they can't be found at intermediate distance.

6- Each electron in the atom has a definite amount of energy depending on the distance between is E.L and nucleus. This energy increases as its radius increases.

7- It was found that the maximum no of energy levels in the heaviest known atoms in their ground state (unexcited) is only seven (K, L, M, N, O, P, Q). Each level has energy expressed by awhole no called principle Q. No.

Ex: The 1st E. Level K its principle Q. no
= 1The 2nd E. level L its principle Q. no
= 2

8- If when atom is excited by heating (Quantum) or by electric discharge the electron will transfer to a higher E. level agrees with the absorbed quantum. The excited electron in the higher E. level is then unstable, so it returns to its original level losing the same quantum of energy, which it gained during excitation in the form of radiation have definite wavelength and frequency.

❖***Remarks:***

1- The quantum: Is the amount of energy gained or lost when an electron jumps from one E. level to another.

2- The difference in energy between levels (Q) is not equal i.e. the difference in this energy decreases further from the nucleus. This means that the quantum of energy required to transfer an electron from one energy level to another is not equal.

3- The electron does not move from its level to another unless the energy absorbed or emitted is equal to the difference in energy between 2 levels i.e. one quantum.
(There is no half quantum for instance). Q can't be divided or doubled
GIVE REASON: It is wrong to say that e' to be transferred from

E.L (K) to E.L (M) needs amount of energy equals 2 quantum.

<u>EXCITED ATOM:</u>

- It is an atom that acquired an amount of energy (Q) sufficient to transfer its e's from their original E.L to higher ones.

Advantages of Bohr:

Bohr's atomic theory succeeded in the following ways:1- It explained hydrogen atom spectrum.

2- He introduced the idea of quantum no to detect energy of electrons in energy levels.

3- He proved that electrons during that electrons during rotation around the nucleus in ground state do not radiate energy, so they will not fall back to the nucleus.
<u>(a reconcilion between Rutherford and Maxwell).</u>

Disadvantages of Bohr's theory:

1 – If failed to explain the spectrum of any other element even that of He except hydrogen (Simplest Electronic System).

2 – He considered the electron as a (-ve) particle only and did not consider that it also has wave properties.

3 – He postulated that it is possible to determine precisely both speed and location of an electron at the same time. This is experimentally
impossible because the apparatus used will change either speedor location of electron so the result won't be accurate.

4 – He described the electron when moving in a circular planer orbit,
which means that hydrogen atom is planer. Later it was confirmed
that hydrogen atom has 3 dimensional co-ordinates.

Bohr's Theory

The atomic spectrum

Studying and explaining the atomic spectrum was the key to he atomic
structure in 1913 and deserved noble prize at 1922.

ATOMIC EMISSION SPECTRUM:

1 – By heating gases or vapours of substances to a high temperature (by heat or electricity) under low pressure it produces light.
2 – By using spectroscope we find that this light consists of a fixednumber of coloured lines called *line spectrum.*

LINE SPECTRUM OF ANY ELEMENT:

By exp. Proved that spectrum line differs from one element to another like finger prints.

N.B:

Line spectrum of sun rays shows that composed of *hydrogen and Helium.*

The principles of Modern Atomic Theory:

1 – Dual nature of electron.

2 – The Heisenberg uncertainty principle.

3 – Finding the mathematical expression which describes the wave motion of electron, its shape and its energy.

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1 - THE DUAL NATURE OF THE ELECTRON

The experimental data showed that the electron has a dual nature i.e.

a) It is a material particle. *b) It has wave properties.*

**** De Broglie principle:***

Every moving body (such as electron or the nucleus of an atom or whole molecule)is associated with (accompanied by) a wave motion (or matter waves) which has some properties of light waves.

The matter wave motion differs from electromagnetic waves n …………

Matter Waves	Electromagnetic Wave
1- They are not separating from the moving bod	1- They are separated from the moving body
2- Their speed is not equal to the speed of light	2-Their speed is equal to the speed of light

2- The Heisenberg uncertainty principle: (quantum mechanics)

It is practically impossible to determine both position and the velocity of
the electron exactly (precisely) at the same time. We can only say that it is probably to a greater r lesser extent to locate the electron in this or in that place. This is to speak in terms of probability seems to be more precise.

3. The wave equation for motion of electron inside the atom: (Schrodinger wave mechanics theory): He applied the ideas ofBlanck,
Einestein, De Broglie and Heisenberg so he shows that:

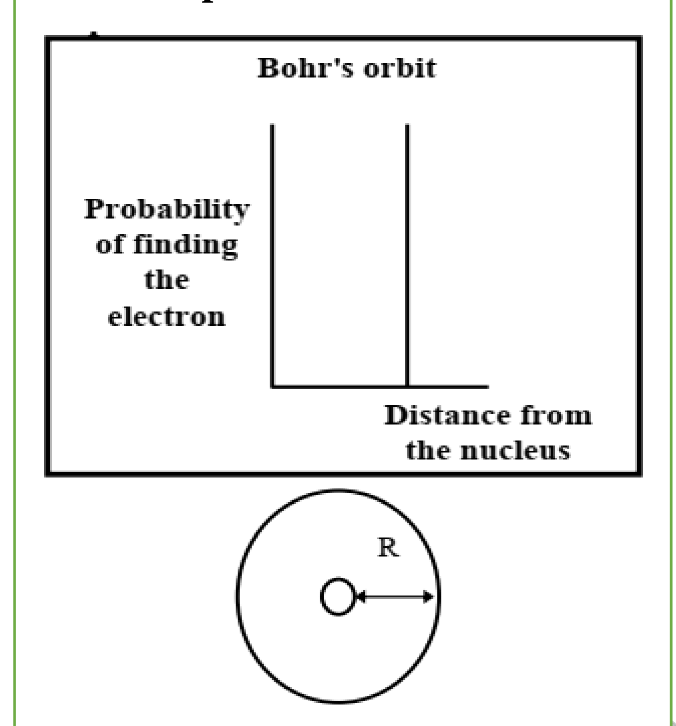
❖ It is possible to determine the allowed energy levels of the electron and define the region of space around the nucleus where it is most probable to find the electron in each energy level.

❖ The electronic motion around nucleus has a wave properties therefore the position to use the term electron cloud to describe any orbital.

Electron Cloud: (used to describe any orbital)
"Area of space around the nucleus where there is a great probability for
finding electrons in all direction and all positions."
❖ The difference between the orbit and orbital concepts according to both Bohr and the wave mechanics theories:

Bohr s theory

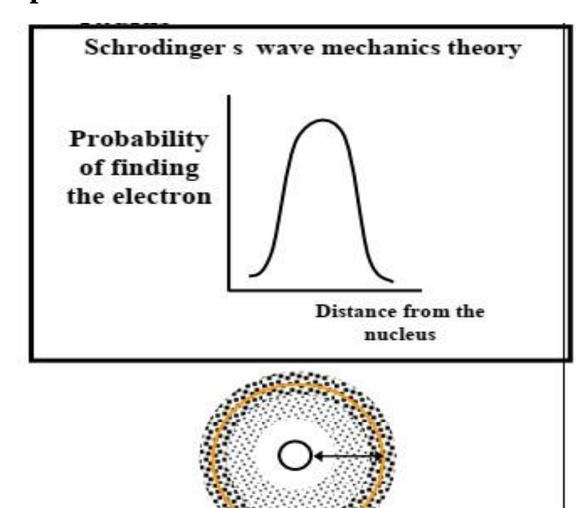
- **It is a circular planer orbit with particular radii**



Schrodinger s wave mechanics

It is an electron cloud used to describe any orbital

- **Electron cloud the regions of high density of dots represent the region of high probability of finding the (e) from which it is possible to define the atomic radius**



❖ The mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers.

Quantum Numbers:

These nos define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they definethe energy, shape and direction of orbitals.

1-Principle Q.no (n).

2-Subsidiary or orbital (azimuthal) Q. no

(L).3- Magnetic Q. no (m).

4- Spin Q. no (ms).

1- Principle Q No (n):

Bohr had used this no to define the following:

1-Order of principle E. levels their number in the heaviestknown atom in the ground state is seven.

2-No of electrons required to fill a given E. level = two timesthe square of the level no (2n2).

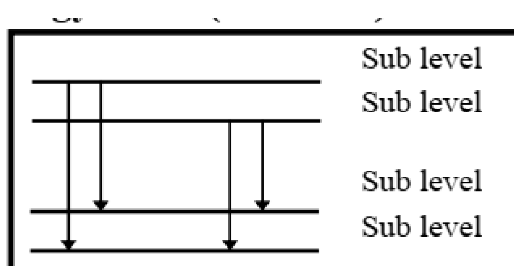
-1st E.L	**K**	**Is filled with**	**2 electrons**
-2nd E.L	**L**	**Is filled with**	**8 electrons**
-3rd E.L	**M**	**Is filled with**	**18 electrons**
-4th E.L	**N**	**Is filled with**	**32 electrons**

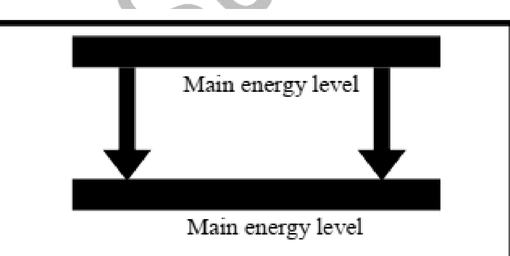
*But this rule does not apply to the last three levels (O, P, Q). However, the atom becomes unstable if no of electrons exceeds 32 electrons on anylevel.

2. ***Subsidiary Q No (L):***

1- Used to detect the no of sub levels in each E. level.

2- The energy sub levels take the symbols s, p, d, f. this is shown by the scientist Somerfield. When he used a spectroscope which has a high resolving power, he found that the single line (which represents electron transition between two different energy levels) is indeed a number of fine spectral lines which represents electron transition between very near energy levels (sublevels).





3- No of sublevels in each energy level = order of principle energylevel (n).

-1st E.L	K	has 1 sub level	1s
-2nd E.L	L	has 2 sub level	2s, 2p
-3rd E.L	M	has 3 sub level	3s, 3p, 3d
-4th E.L	N	has 4 sub level	4s, 4p, 4d, 4f

N.B:

- **Energy of sub levels of same E. level is not equal.**

f>d>p>S

- **Energy of same sublevels but of different E. levels also differ in energy.**

<u>Ex:</u> **4d>3d** **4p>3p>2p**

• There is a small difference in the energy between sub-levels.

3. Magnetic Q No (m):

Detected by Zieman when he exposed spectral line to strong magnetic field, he found that each line divides into many lines, so he concluded that each E. sublevel has no orbitals.

- Magnetic Q No is characterized by:

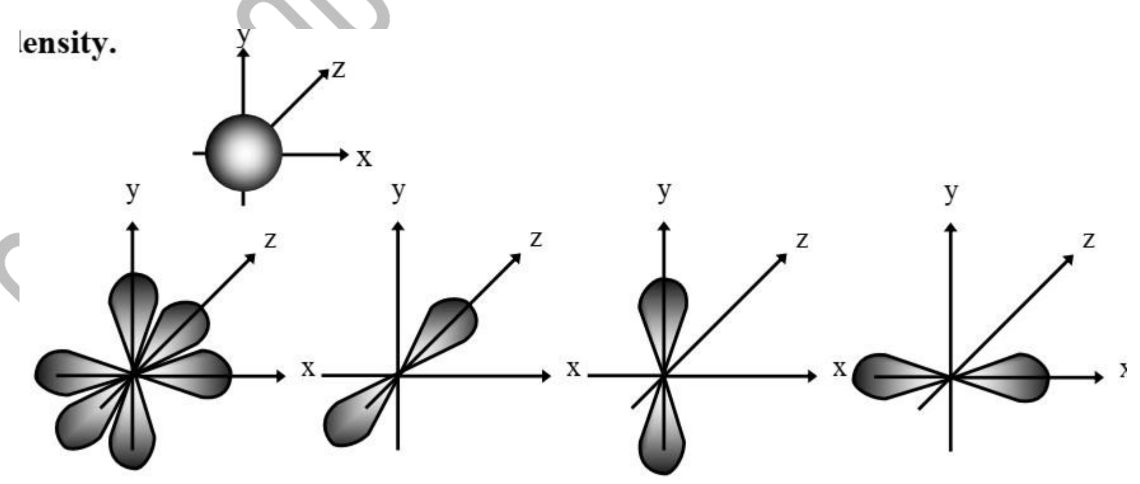
1- Used to detect no of orbitals in each E. sub level and their direction in
space.

2- Sublevel (S) has one orbital of spherical symmetrical

shape.3- Sublevel(P) has 3 orbitals.

• Each orbital (Px,Py,Pz) is perpendicular to the other two.

• Also the electron cloud of each orbital takes the form of 2 pears meeting head to head (dumb – bell shaped) at a node i.e. point of zero electron density.



4 -Sublevel (d) has 5 orbitals**.**

5 -Sublevel (f) has 7 orbitals**.**

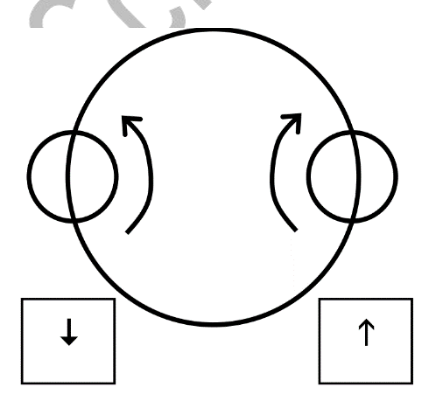
6 **-**Orbitals of the same sub level are equal in energy and shape**.**

Ex: px = py = pz

7 -No of orbitals in each E. level = square order of E. level = n2**.**

4 -SPIN Q. NO (MS):

- *Detects the direction in which the electron spins around its axis during its rotation around the nucleus***.**
- *Each orbital can be saturated by 2 electrons, one electron spins around its axis clockwise while the other electron spins anti – clockwise in order to from 2 opposite magnetic fields to decrease the force of repulsion between them which keep the atom stable***.**
- *It has only two possible values + 1/2 – 1/2*



GIVE REASON*: Each orbital carries 2 electrons although they are negatively charged.*

Summary of the relationship between the principle E.L, sub levels orbitals and no of electrons:

1 No of energy sublevels = order of principle level (n).

2 No of orbitals within a principle level = square the no of thelevel (n^2).

3 No of electrons occupying a given E. level = two times thesquare order of this

level ($2n^2$)

Ex:

Order of E.L.	E.L	No of sub-levels (n)	No of orbitals (n2)	No of electrons (2n2)
1st	K	1 (1s)	1	2
2nd	L	2 (2s, 2p)	4	8
3rd	M	3 (3s, 3p, 3d)	9	18
4th	N	4 (4s, 4p, 4d, 4f)	16	32

-PRINCIPLE OF DISTRIBUTING ELECTRONS:

- **Aufbau (building – up) principle:**

"Electrons occupy energy sublevels in an ascending order according to increasing energy where the lowest energy sublevel is filled 1st".

Ex: 4s is filled before 3d as energy of 4s < 3d.

K	1S			
L	2S			2p
M	3S			3p
N	4S		3d	4p
O	5S		4d	5p
P	6S	4F	5d	6p
Q	7S	5F	6d	

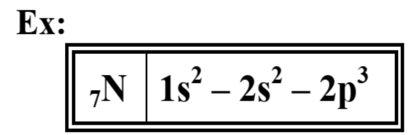
Examples:

Na 11	$1s^2$	$2s^2$	$2p^6$	$3s^1$				
Ca 20	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$		
Zn 30	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	

Hund's rule:

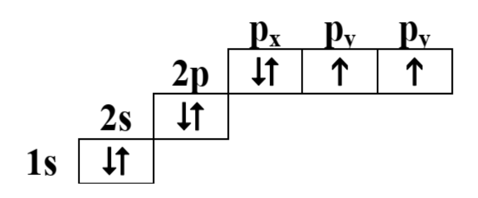
State that: "No electron pairing takes place in a given sublevel until each orbital contains one electron."

Ex:



$_8O$	$1s^2 – 2s^2 – 2p^4$

			p_x	p_y	p_y
		2p	↑	↑	↑
	2s	⇅			
1s	⇅				



Atom is stable when the outer sub-level is half completely filledwith e's.

3 Remarks:-

1. Electrons are preferred to be unpaired before pairing because according to Hund's rule on pairing electrons in the same orbitals, they will repel decreasing stabilityof the atom.
2. Electrons prefer to be paired with another electron than to transfer to a higher sub-level, as the energy needed to transfer it to a higher sub-level.
3. Also the spin of single electrons must be in the same direction because this gives the atom more stability.

Another E. configurations (to the nearest noble gas)

<u>Ex:</u>

Li 3: (He_2) ,$2S_2$

Na 11 : $(Ne_{10}$) , $3S_1$

K19 : (Ar_{18}), $4S_1$

Questions

- Choose:

1- 1^{st} scientist defines the element is ……………

a) Dalton b) Rutherford c) Boyl d) Thomson

2- Substance composed of 4 components which are (water, air, dust and fire) was the idea of ……….

a) Bohr b) Rutherford c) Dalton d) Aristotle

3- To prove that cathode rays taking part in all substances ……………

a) Have thermal effect.
b) Transfer in straight line.
c) Have tiny porticles.
d) Don't changing substance of cathode or kind of gas.

4- On heating gases or vapours under low pressure at high temperature

a) Absorbs light. b) Gives light.
b) Gives gamma rays. D) Gives alpha rays.

Chapter two

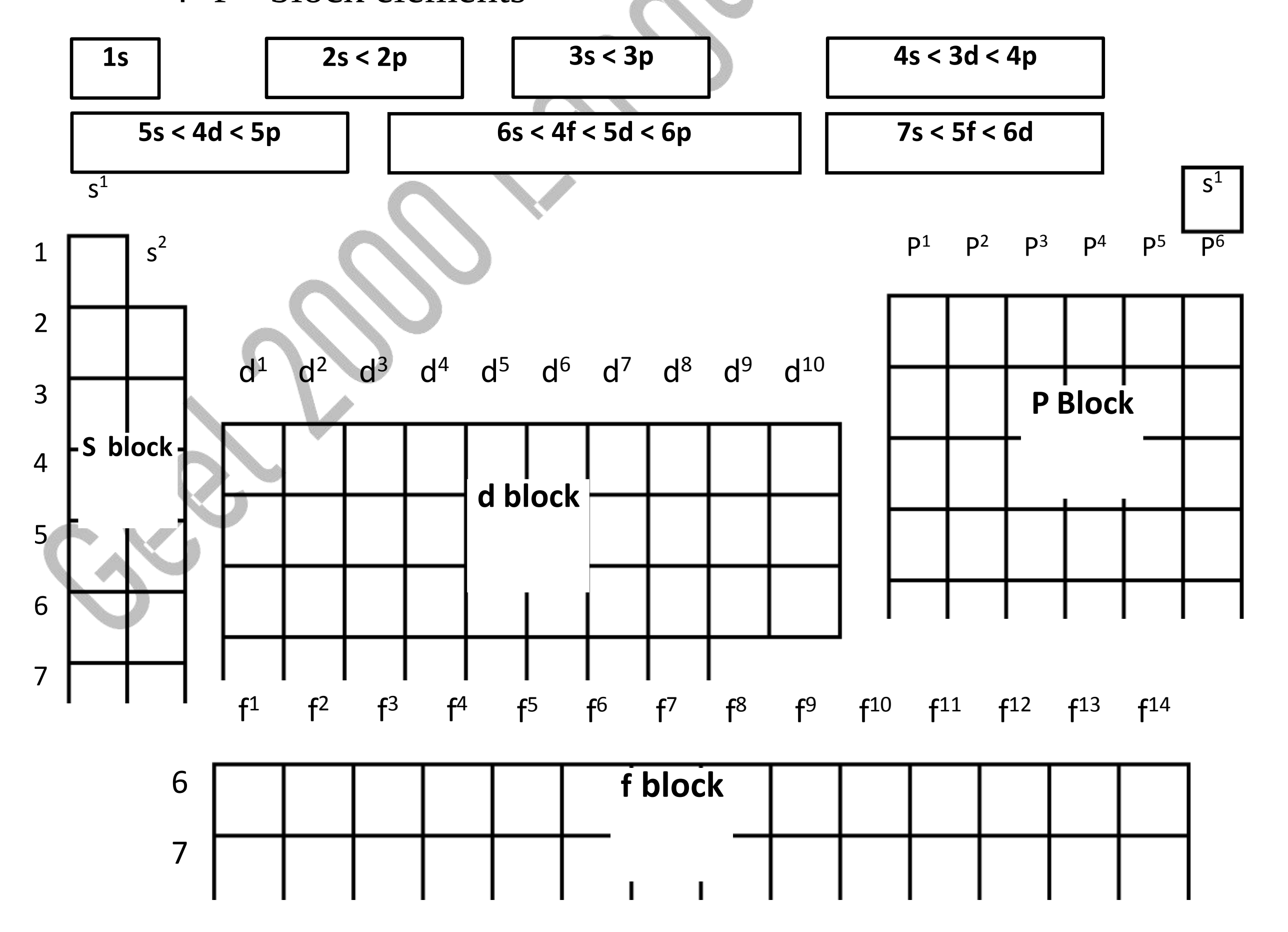
Classification of elements and the long periodic table

The long form periodic table:
It depends on the building up principle (Auf - bau) The elements are arranged in an ascending order according to their atomic number (each element have one electron more than the element before it) .

The long form periodic table : the elements were arranged in an ascending order according to their atomic numbers

The periodic table is divided into four main blocks

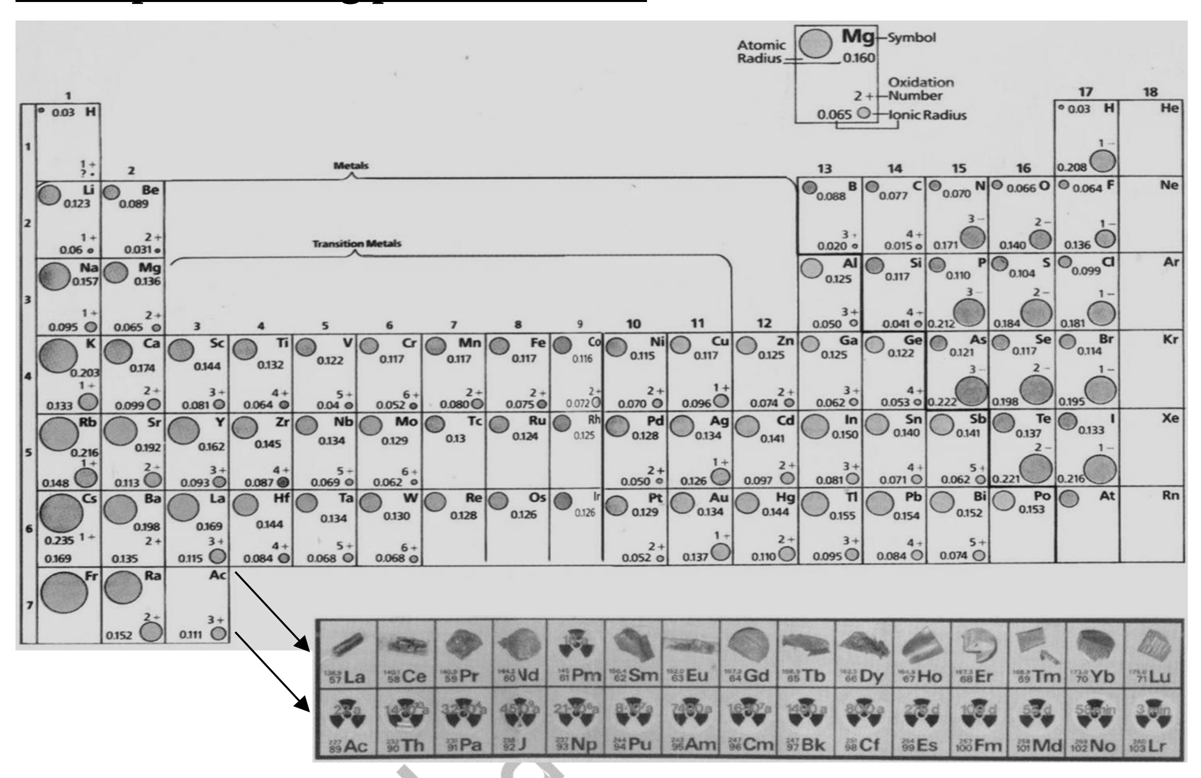
1- S – block elements
2- P – block elements
3- d – block elements
4- f – block elements



Sequence of energy sublevels

1S < 2S < 2P < 3S < 3P < 4S < 3d < 4P < 5S < 4d < 5P < 6S < 4f < 5d < 6P < 7S < 5f < 6d

Description of long periodic table :



1- S – block consists of two groups because the S – sublevel consists of one orbital which is filled with two electrons only .

2- P – block consists of (6) groups because the P – sublevel consists of three orbitals which filled with six electrons .

3- d – block consists of (10) groups because the d – sublevel consists of five orbitals which are filled with ten electrons .

4- f block are separated from the table so that the table is not too wide (long)

5- The first period contains two elements because it consists of elements of the sublevel 1S = 2 electrons .

6- The second period contains eight elements because it consists of the sublevel (2S + 2P) = 8 electrons .
7- The third period contains eight elements because it consists of element of sublevel (3S + 3P) = (2 + 6) = 8 electrons .

8- The fourth period contains eighteen elements because it consists of elements of the sublevels (4S + 3d + 4P) = 2 + 10 + 6 = 18 electrons .

9- The fifth period contains (32) elements because it consists of elements of the sublevels (5S + 4d + 5P) = 2 + 10 + 6 = 18 electrons .

10- The six period contains (32) elements because it consists of elements of the sublevels (6S + 4P + 5d + 6P) = 2 + 14 + 10 + 6 = 32 electrons .

<u>How can you find the location and the type of element in the periodic table ?</u>
1- Write the electronic configuration of element in quantum levels .
2- Number of period = The maximum value of principle energy level
(quantum number) .

<u>Example :</u>
Find the number of period and group for each of the following element :
Na : The atomic number = 11
Cl : The aomic number = 17

<u>Noble gases</u> : They are the elements of the last column of the P – block all their energy levels are completely filled with electrons .

<u>The representative elements :</u> They are the elements of main group (S – and P – blocks) all their energy levels are completely filled with electrons except for the external energy level .

The transition elements : They are the elements of the d – block all their energy levels are completely filled with electrons except for the two external energy levels .

The inner transition elements : They are the elements of the f – block all their energy levels are completely filled with electrons except for the three external levels .

Trends and periodicity of properties in the periodic table

The atomic radius :
We cannot determine the atomic radius because electron has a wave motion so it is impossible to determine exactly the location of an electron around the nucleus .

The atomic radius :
It is half the distance between centers of two similar atoms in a diatomic molecule .

The bond length : It is the distance between the nuclei of two bonded atoms .

There are many methods to measure the bond length such as :

1- X – ray . 2- Electron diffraction .

Examples

1- The bond length in the chloride molecule Cl – Cl is 1.98 Å and the length between carbon and chloride atoms C – Cl is 1.76 Å . Calculate the atomic radius of carbon .

<u>solution</u>

The atomic radius of chlorine = $\frac{1.98}{}$ = 0.99 Å

The atomic radius of carbon = 1.76 – 0.99 = 0.77 Å

1- **The bond length in the molecule of NH_3 is 1.0 Å and the bond length in the molecule of H_2 is 0.6 Å . Calculate the bond length in nitrogen molecule (N_2) ?**

<u>solution</u>

The atomic radius of hydrogen = $\frac{0.6}{}$ = 0.3 Å

The atomic radius of nitrogen = 1 – 0.3 = 0.7 Å

The bond length of the N_2 = 0.7 × 2 = 1.4 Å

<u>Atomic radius decrease in period by increasing atomic number ?</u>

Because of increase in the atomic number gradually makes to increase the positive nuclear charge therefore the attractive force of the nucleus for electrons will be increased and atomic radius decreased .

<u>Atomic radius increases in group by increasing of atomic number why ?</u>

Because the increase in atomic number makes to increase the number of energy level screen the attractive force of the nucleus

for valence electrons and increase the repulsion force between electrons therefore the atomic radius increases .

N.B.

1- The cation (+ve ion) radius is smaller than that of its atom ?

This is due to the increasing of positive charges of the protons which attract the valency electrons leading to a decrease in the cation radius .

2- The anion's radius is bigger than that of its atom ?

This is due to the increase in the number of the negative charges in the shells therefore the repulsive force between the electrons increases so the shells move a part and this leads to increasing of the anionic radius than that of its atom .

Ionization potential

It is the amount of energy required to remove the smallest bounded electron completely from an isolated gaseous atom

$M \longrightarrow M^+ + e^-$ ΔH = + ve KJ/Mole

The atom losses electrons and converted into positive ions . It has a positive value (ΔH = + ve KJ/Mole) .

1- The first ionization energy : It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge .

$M \longrightarrow M^+ + e^-$ ΔH = + ve KJ/Mole

2- The first ionization energy of noble gas is very high ?

Due to the stability of their electronic configuration because it is difficult to remove an electron from completely filled shell .

3- The ionization energy of element of group (5A)

(N_7 – P_{15} – As_{33} – Sb_{51} – Bi_{83}) is much greater than any element have the same period because the outer most energy sublevel (P) has three electrons and it is half filled with electrons (nP) and this gives the atom of the element some extra stability so the ionization energy is greater .

4- **<u>The ionization energy of sodium is much smaller than that of chlorine ?</u>**

Because the atomic size of chlorine is smaller than that of sodium so the attractive force of the nucleus on the valence electrons in the case of chlorine is more strongly and the electrons valence need a higher energy to be separated from the atom .

5- **<u>Ionization energy increases period ?</u>**

Because the positive nuclear charge gradually increases with the increase of the atomic number led to decrease the atomic radius and increase the attractive force of the nucleus on the valence electrons therefore the electrons needed m large high energy to remove (separated) from the atom .

6- **<u>Ionization energy decreases in group ?</u>**

Due to the increase in the atomic size and screen of the attraction force of the nucleus on the valence electrons therefore the electrons needed a smaller value of energy to separated from the atom .

7- **<u>The ionization energy of elements of group (2A) is much greater than any element have the same period ?</u>**

Because the outer most energy level (S) of the element of group (2A) is completely filled with electron (nS^2) and this gives the atom of the element some extra stability so the ionization energy is much greater .

Electron affinity

It is the amount of energy released when an extra electron is added to a neutral gaseous atom to form an anion (- ve ion)

$X + e^- \longrightarrow X^- + \text{energy}\ \Delta H = -\text{ve KJ/Mole}$

***In the horizontal periods electron affinity increases with the increase in atomic number ?**

Due to the atomic radius (size) gradual decrease so it becomes easier for the nucleus to attract the new electron .

***The electron affinity decreases in group ?**

Due to the increase of the atomic volume with increase atomic number and this leads to the screening of the attraction force of thenucleus on the valence electrons .

Exception cases :

Beryllium has a relatively high of electrons affinity due to the stability of its atom that has completely filled orbitals ($1S^2$, $2S^2$) ?

Because the outer most energy sublevel (nS) is completely filled with electron and it gives the atom some extra stability .

Elements of the fifth group (N_7 , P_{15}) have a lower value of electron affinity

Because the outer most energy sublevel (nP) has three electrons and it is half filled with electrons it gives the atom some extra stability (N_7 : $1S^2$, $2S^2$, $2P^2$).

Noble gases have not (small) electron affinity

Because all energy sublevels are completely filled with electrons which gives the atoms great stability .

Electron affinity of Fluorine (F_9) is less than that of chlorine (Cl_{17}) ?

Because the atomic radius (size) of fluorine atom is smaller than that of chlorine atom and when fluorine atom gains electron it is

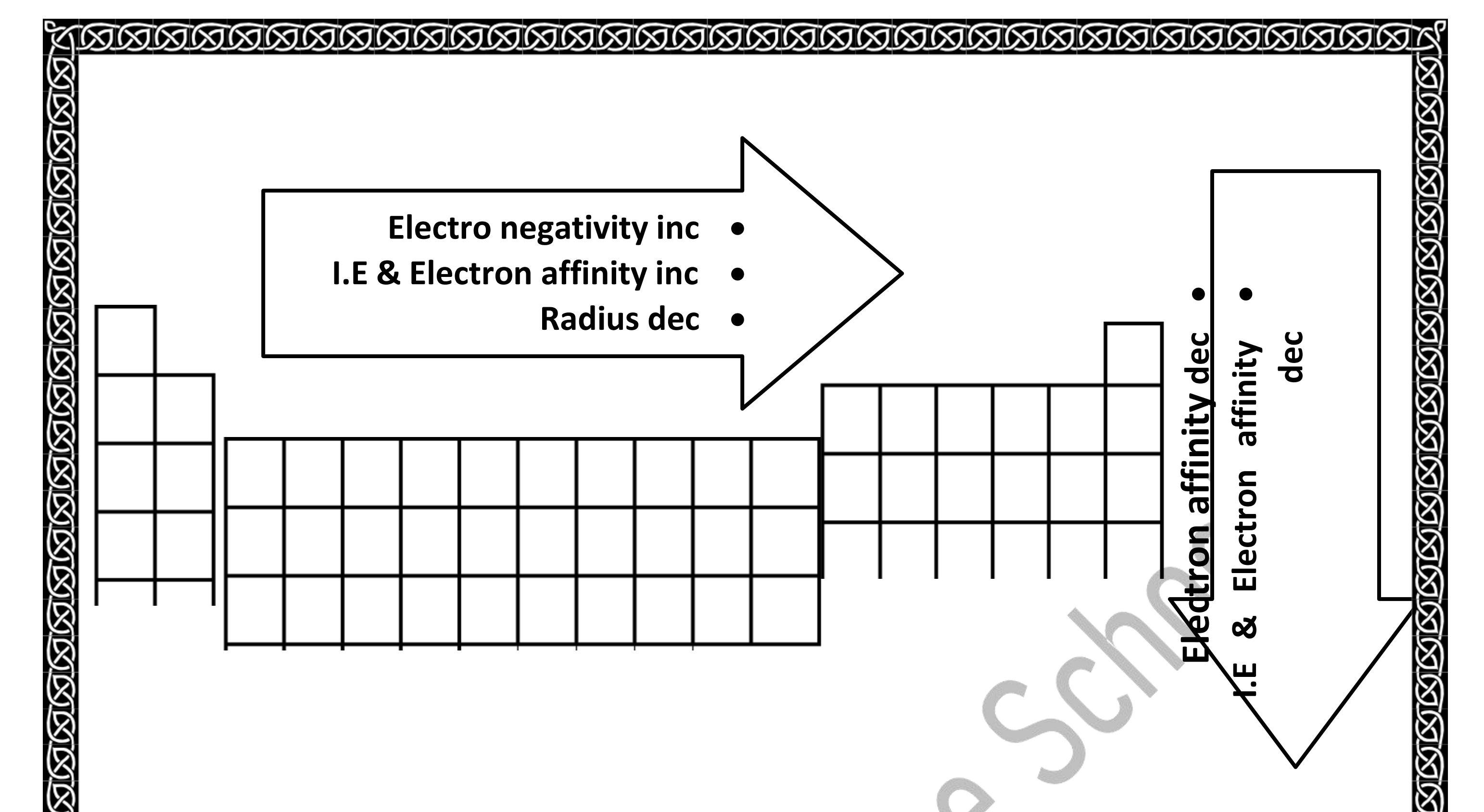
affected by a great repulsion force bigger than that in chlorine atom and fluorine atom is very small size .

Electro negativity

The tendency of an atom to attract the electrons of chemical bond to itself .
It is the querage of the ionization potential and electron affinity .

Compare between ionization energy , electron affinity and electro negativity

Ionization energy	Electron affinity	Electro negativity
It is the amount of energy needed to remove the least connected electron bond in a single atom	It is the amount of energy released when an extra electron is added to a neutral single atom to form an ion	It is tendency of an atom to attract the electrons of chemical bond to itself
It refers to the atom in its single state	It refers to the atom in its single state	It refers to the atoms which linked together in the molecule
It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius
The atom losses electrons and converted into positive ion	The atom gains electrons and convers into negative ion	
It has a positive value	It has a negative value ΔH = - value type of reaction is exothermic	



Differentiation between metal and non-metal

Metals	Non metals
Valence shell has less than help its capacity of electrons (1 or 2 or 3)	Valence shell has more than half its capacity of electrons (5 or 6 or 7)
They one called electropositive elements .	They are called electronegative elements
They have Relatively large atomic radius therefore the ionization energy and electron affinity and electro negativity have smell values	They have small atomic radius ionization energy and electro negativity and electron affinity have high value
They have a good conductivity of electricity .	they have a bad conductivity of electricity .

N.B.

1- Fluorine is the strongest non metal while caesium is the strongest metals Because fluorine has smallest radius while caesium has the biggest radius .

2- Metals are considered as electropositive elements because metals lose electrons to form positive ions .

3- Non metals are considered as electronegative elements because non metals gains electrons to from negative ions .
4- Metal are good conductors of electricity because they have few valence electrons which can transfer easily from one position to another in the metal structure .

5- Non metals are bad conductors of electricity because their valence electrons are strongly bounded to the nucleus due to the small atomic size therefore it is difficult for the valency electrons to be transferred .

6- Metals have small values for ionization energy and electron affinity because they have large atomic radius .

7- The strongest metal lie at the bottom on the left and side of the periodic table because of the increasing in the atomic number the atomic radius increases gradually so the attractive force of the nucleus to the valency electrons decreases therefore it is very easy for the atom to lose the valency electrons so the metalic property increases .

Metalloids

1- They are elements whose valency shell contains 4 electrons .
2- Sometimes they act as metal (when they gains electrons) .
3- They act as semiconductors (Boron – silicon) which are used in or transistors and knows as semiconductors electronic instruments .

Decreases → increase

metals → metalloids → non metal

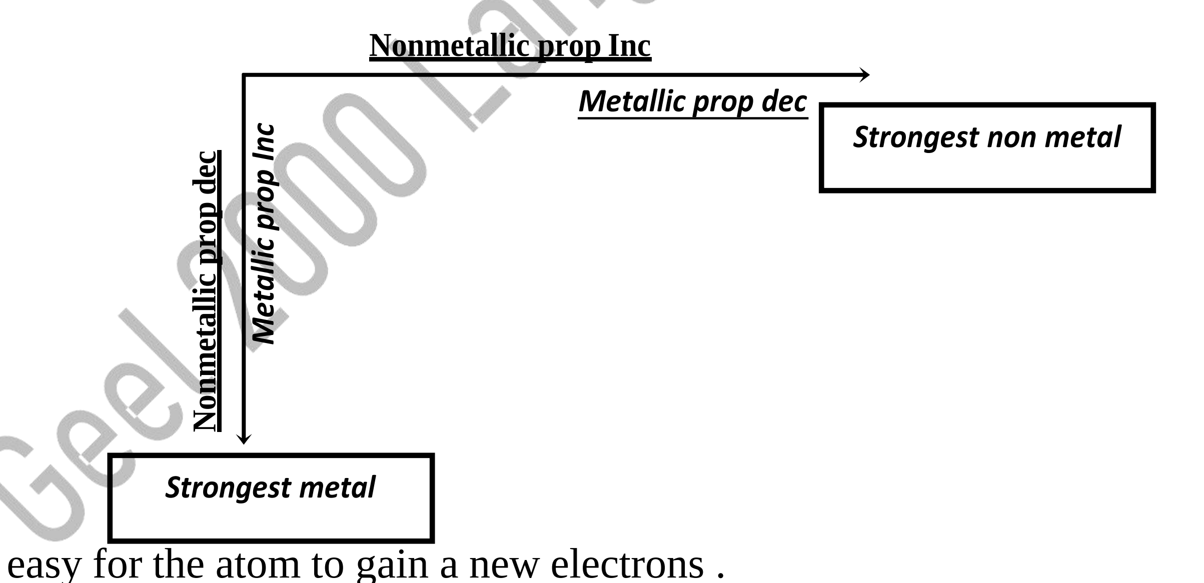
* Because the atomic size decrease and the attractive force of nucleus on the valence electron will be increased therefore it is difficult for the atom to lose the valence electron .

Metalloids increase in period because the atomic size decrease and the attractive force of nucleus to electron will be increased therefore it becomes easier for the nucleus to gain a new electron .

In groups : Metals increase in group because the increase in the atomic number makes to increase of energy levels and screen (decrease) the attractive force of the nucleus on the valence electrons therefore it is easy for the atom to lose the valence electrons .

In groups : non metals decrease because the increase in the atomic number makes to increase the energy level and screen the attractive force of nucleus on the valence electrons therefore it is difficult for the atom to gain a new electron .

Fluorine is considered as the strongest non – metal because the atomic size of fluorine is very small therefore the attractive force of the nucleus to the electrons will be increased therefore it is very



easy for the atom to gain a new electrons .

Acidic – basic properties

Acidic oxides	Basic oxides
They are non – metalic oxides such as CO_2 , SO_2 , SO_3 , P_2O_5	They are metalic oxides such as Na_2O , K_2O , MnO , CaO , BaO
They dissolve in water to form acids $Co_2 + H_2O \longrightarrow H_2CO_3$ $SO_2 + H_2O \longrightarrow H_2SO_3$ $SO_3 + H_2O \longrightarrow H_2SO_4$ $P_2O_5 + H_2O \longrightarrow 2H_3PO_4$	Some basic oxides dissolve in water to form alkalies and others are not $Na_2O + H_2O \longrightarrow 2NaOH$ $K_2O + H_2O \longrightarrow 2KOH$ $CaO + H_2O \longrightarrow Ca(OH)_2$ $MgO + H_2O \longrightarrow Mg(OH)_2$
They react with alkalis to form salt and water $CO_2 + 2NaOH \longrightarrow Na_2CO_3 + H_2O$ $SO_2 + 2NaOH \longrightarrow Na_2SO_4 + H_2O$	They react with acids to form salt and water $Na_2O + 2Hcl \longrightarrow H_2O + 2Nacl$ $MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
They do not reaction with Acids	They do not react with alkalis

The oxyacids : are acids that contain hydrogen, oxygen and a third element usually a non-metal .

- It can take the following symbols: - M On (OH)m

where , (M) is the atom of the element .

(n) is the number of oxygen atoms .

(m) is the number of hydroxyl groups .

The strength of oxyacids are depends on the number of oxygen atoms which does not linked with hydrogen atoms when this number increase , the strength of the acid .

Acid Name	The number of free atoms of Oxygen	Acidic Property
H_4SiO_4	-	Weak acid
H_3PO_4	1	Moderate acid
H_2SO_4	2	Strong acid

$HClO_4$	**3**	**Very strong acid**

Amphoteric oxides :

1- They react with bases and acids to form (give) salt and water in both cases .

2- They have both acidic and basic properties such as Al_2O_3 , ZnO , Sb_2O_3 , SnO .

How can you obtain CuO from mixture of (CuO + Al_2O_3) ?

3- KOH is strong base than of NaOH because the atomic radius of K is greater than that of Na so the bond strength between K and hydroxide group (OH^-) decreases therefore it is easy to separate hydroxide ions to form strong base .

4- HF is weak acid because the atomic size of fluorine atom is very small so it has ability to attract the electrons bond to itself therefore it is difficult to separate hydrogen ions to form strong acid .

5- HI is stronger acid because the atomic (size) radius of iodine atom is largest to it has a weak ability to attract the electron bond to itself therefore it is easy to separate hydrogen (H) ions to form strong acid .

Oxidation number : at last we defined it as it is the number of hydrogen atoms that combine with or can be replaced by an atom of the element .

The modern definition : the number of single (unpaired) electrons in the valence shell of the atom .

N_7 $1S^2$ $2S^2$ $2p^3$ Trivalet
O_8 $1S^2$ $2S^2$ $2p^4$ Divalent
F_9 $1S^2$ $2S^2$ $2p^5$ Mono valet

Oxidation number : it is a number that refers to the electric charge (+ve or -ve) that atom would have in the compound .

Rules for assigning oxidation numbers

Oxidation number of Oxygen = 2- in most of its compounds except in peroxides e.g. (hydrogen peroxide H_2O_2) , sodium peroxide Na_2O_2 and potassium peroxide K_2O_2 is (1-) supper oxide e.g. KO_2 the oxidation number of oxygen is (½-)

The oxidation number of oxygen = 2+ in OF_2 because the electro negativity of fluorine is higher than the electron negativity of oxygen .

The oxidation number of hydrogen in its hydrides = 1- such as NaH , KH , CaH_2 .

The oxidation number of any chemical compound = zero because the algebraic sum of the oxidation number of its atom = zero .

Such as NaCl , HCl , CuSO4

The oxidation number of any element in its pure state = zero .
e.g. Fe , Cl , Na , O_2S , P4 .

The oxidation number for any atomic group (poly atomic ion) = the number of charge on the group (ion) such as OH^- , $SO4^{--}$, $CO3^{--}$, $NH4^+$, $PO4^{---}$.

Examples

Calculate the oxidation number of chlorine in these compound , and for sulphure

Nacl **Nacl = zero** **+1 +Cl = 0** **Cl = -1**	**$NaClO_2$** **$NaClO_2$ = zero** **1+Cl 2-×2 = zero** **Cl = +3**	**$NaClO_4$** **$NaClO_4$= zero** **+1 Cl -8 = 0** **Cl = +7**	**H_2S = zero** **2×1+S= zero** **S = 2-**	**$SO3^{--}$** **SO3 = -2** **S -6 = -2** **S = 4**
K_2S **K_2S = zero** **C + S = 0** **S = -2**	**$Na_2S_2O_3$** **$Na_2S_2O_3$ = zero** **2×1+S_2 -6= zero** **S = 2**	**S = zero**	**H_2SO_4= zero** **2 + S-8= 0** **S = 6**	**$SO4^{--}$** **SO4 = -2** **S -8 = -2** **S = +6**

<u>The oxidation process</u> : it is the process of losing electrons due to increase the oxidation number for the element.

<u>The Reduction process</u> : it is the process of gaining electrons due to decrease the oxidation number for the element .

Oxidation	**Reduction**
Zero ⟶ (+) value (+) small value⟶ (+) big value (-) value⟶ zero (-) big value⟶ (-) small value	(+) value ⟶ zero (+) big value⟶ (+) small value zero⟶ (-) value (-) small value⟶ (-) big value

Explain the type of change (oxidation or Reduction for carbon in this Reaction .

C_2H_5OH ⟶ CH_3CHO ⟶ CH_3COOH ⟶ CO_2

C = -2 ⟶ C = +1 ⟶ C = zero ⟶ C = +4

Oxidation Oxidation Oxidation

Explain the type of change (oxidation or reduction) that of chromium and iron in this reaction .

$K_2CR_2O_7 + 6FeCl_2 + 14HCl \longrightarrow 2KCl + CrCl_3 + 6FeCl_3 + 7\ H_2O$

Solution

$K_2Cr_2O_7 \xrightarrow{\text{Reduction}} CrCl_3$ **$FeCl_2 \xrightarrow{\text{Reduction}} FeCl_3$**

$Cr = +6 \longrightarrow Cr = +3$ **$Fe = +2 \longrightarrow Fe = +3$**

R

Gradation of oxidation number in periodic table :

Group	I	II	III	I V	V	V I	V II
Element	Na	Mg	Al	Si	P	S	Cl
Oxide Oxidation Number	Na_2O 1 +	MgO + 2	Al_2O_3 + 3	SiO_2 + 4	P_2O_5 + 5	SO_3 + 6	Cl_2O_7 + 7

In the case of hydride

- The oxidation number of element (metal) of 1st and 3rd groups in their compounds agrees with group number (it takes positive oxidation number to which they belong)
- Most elements in the middle of the table have variable oxidation number in different compounds .

The oxidation number of the elements in groups from IV to VII = the group number -8 = **-** value because the elements gain electrons to complete a full shell .

Group	**I**	**II**	**III**	**I V**	**V**	**V I**	**V II**
Element	Li	Be	B	C	N	O	F
Hydride Oxidation Number	LiH 1 +	BeH_2 + 2	BH_3 + 3	CH_4 + 4	NH_3 - 3	H_2O - 2	HF - 1

The oxidation number of noble gases of (group zero) is zero because they do not combine to form compounds .

Questions on Chapter 2

<u>Choose the best answer:</u>

1- The sublevel (d) contains .... orbitals, it can accommodate up to .... electron . (2 - 3 - 5 - 10 - 14)

2- An element of atomic number 10, the number of sublevels filled with electrons are. .... (2 - 3 - 5)

3- The valency shell of the atoms of fifth group elements contain...........

a) np4 b) np5 c) np d) np3

4- Group 5-A elements are classified as .... block. (S-P-d-f)

5- Sublevel (f) saturated with number of electrons equals..... .

a) 6 b) 16 c) 14 d) 32

6- The fourth energy level is saturated by a number of electrons equals .... a) 8 b) 16 c) 32 d) 64

7- Electron affinity in group decrease...........

a) Up on increasing atomic number only

b) up on decreasing atomic number only

c) up on increasing atomic volume only

d) a and c

8- The oxidation number of hydrogen in hydrogen molecule is (1, 2, 3, 0) and its oxidation number of hydrogen chloride is (1, 2, 3, 0).

9- The oxidation number of the first group elements equal....

a)+2 b)+l c)-2 d)-l

10- The oxidation number of oxygen in H2O2 is...........

a)+2 b)+l c)-2 d)-l

11- The atomic number of element is 29, so, its electronic structure is..........

a) (Ar) $3d^{10}$, $4S^{1}$

b) (Kr) $4P^{4}$, $5S^{2}$

c) (Xe) $6P^{6}$, $6S^{2}$

12- The electron affinity of Fluorine as compared to chlorine is ....

a) Less b) more c) the same d) non of the above

13- The oxidation number of Sulphur in (SO_4) is ....

a) 2+ b)4+ c)2- d) 6+

14- The metallic property increase in the group elements by...........

a) increasing of atomic number only .
b) increasing of ionic potential .
c) increasing of boiling degree .
d) when the radius decrease .

15- The element whose atomic number is 24, its correct electronic configuration is ....

a) (Arlg) 4S2, 3d" b) (Arlg) 4S\ 3d'
c) (Arlg) 4S°, 3d4 d) (Arlg) 4S2, 3d'

16- The element 86X, its electronic configuration in the outer most energy level is ....

a) $7S^1$ b) $3d^5$ c) $4S^2, 4P^5$ d) $6S^2, 6p^6$

17- The element X, its atomic number is 73 is considered from .... series.

a) Lanthanides b) Main transition elements
c) Representative elements d) actinide

18- $CO_2$is............oxide.

a) Amphoteric b) Acidic c) Neutral d) Basic

19- Na_2O is .... oxide .

a) Amphoteric b) Acidic c) Neutral d) Basic

20- $A1_2\ O_3$ is........... oxide.

a) Amphoteric b) Acidic c) Neutral d) Basic

21- The element has the outer most electronic configuration 75 is found in ………

a) the first period, group 7 b) the six period, group 1

c) the seventh period, group 1 d) the six period, group 7

22- The spin quantum number (Ms) decide ....

a) Energy sublevel and their number

b) Number of electrons in each principle energy level

c) Direction of electron movement

23- The weakest acid is ....

a) HC1 b) HBr c) HF

24- Orbitals of the same energy sublevel are ....

a) different in energy but similar in shape

b) similar in energy but different in shape

c) similar in energy and shape

25- The oxidation number of oxygen in OF_2 is ....

a) -2 b) -1 c) +2

26- The element having the electronic configuration (Ar) $3d^5$ $4S^1$ is...........

a) Representative b) Inner transition c) Main transition

27- The maximum number of electrons that saturate, the energy level (n) is ....

a) 2n b) 2n2 c) 3n2

28- Elements of group III are of the block ....

a)S b)P c)d d)F

29- In Mosely periodic table, the elements are rearranged according to:

a) mass numbers b) atomic numbers

c)atomic weight d)a & b

30- The radius of the atom is :

a) the distance between the nucleus and the farthest electron in the atom,

b) half the distance between two combining atoms in a diatomic molecule,

c) the distance between the centers of two similar atoms in a diatomic molecule

d) half the distance between centers of two similar atoms in a diatomic molecule

31- In the same group as we go downwards the ionization energy:

a) increases b)decreases c) not change

32- The electron affinity of fluorine as compared to chlorine is:

a) less b)more c) the same d) a & b

33- Metal ions are usually :

a) neutral b) positive c) negative d) acidic

- While nonmetal ions are usually:

a) positive b) negative c) neutral d) basic

34- The element where the outermost shell contain more than 4 electrons are :

a) metals b) nonmetals c)metalloids d) inert gas

35- The small radius of nonmetal make the value of ionization energy and electron affinity:

a) small b) great c) moderate d) no change

36- Na_2O is :

a) acidic oxide b) basic oxide c) amphoteric d) neutral

37- $A1_2O_3$ is:

a) acidic oxide b) basic oxide c) amphoteric d) neutral

38- P_2O_5 is:

a) acidic oxide b) basic oxide c) amphoteric d) neutral

39- The perchloric acid $HClO_4$ as compared to the suphuric acid H_2SO_4 is:

a) stronger b) weaker c) the same strength

40- The oxidation number of oxygen in potassium oxids K_2O is :

a) (1+) b) (2-) c) (1-) d) (2+)

41- The element which is found in the second period and has many oxidation numbers (5+ ——>3-):

a) nitrogen b) carbon c) aluminium d) sulphur

42- The element which has(1+)as oxidation number and sometimes (1-) is:

a) lithium b) hydrogen c) chlorine d) cesium

43- The element which has the electronic configuration IS^2 $2S^2$ $2P^6$ is:

a) sodium b) neon c) iron d) magnesium

and it is in the (2^{nd} , 3^{rd} , 4^{th}) period .

44- The oxidation no. of chlorine in $HClO_4$ is :

a) (1-) b) (1+) c) (5+) d) (7+)

45- The oxidation no. of manganese in $KMnO_4$ is:

a) (1-) b) (4+) c) (6+) d) (7+)

46- The oxidation on. of oxygen in H_2O_2 is:

a) (2-) b) (1-) c) (1+)

47- The oxidation on. of nitrogen in NH_2OH is:

a) (3-) b) (2-) c) (1-) d) (1+)

48- The oxidation on. of hydrogen in MgH_2 is:

a) (1+) b) (1-) c) (2+) d) (2-)

49- The value of ionization energy in the period increases as the: series

a) atomic radius decreases b) atomic ardius increases

c) atomic number increases d) a & c

50- The element (X) its atomic number is 13 is considered

a) lanthanide's b) main transition element

c) representative element d) actinide

51- The 4^{th} period in the periodic table contains elements

a) 8 b) 18 c) 16 d) 32

52- The 6^{th} period in the periodic table contains elements,

a) 16 b) 32 c) 48 d) 30

53- Electronegativity increases across the periods

a) atomic radius increases b) atomic radius decreases

c) atomic number decreases d) a & c

54- Electron affinity increases across the period as :

a) atomic number increaes b) atomic number decreases

c) atomic radius decreases d) a & c

55- The ionization energy of magnesium I .. than that of chlorine

a) more b) less c) equal